Zumdahl

Chapter 15.1 - 15.3 (Pg 708 - 724)

Common Ion Effect: Add the salt (that has the conjugate base) into a weak acid.

1.0M HF solution

HF 1.0M

$HF \rightleftharpoons H^+ + F^{1-}$ $K_a = 7.2 \times 10^{-4}$

$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^{1-}$ $K_w = 1.0 \times 10^{-14}$

$$K_a = \frac{[H^+]_E[F^{1-}]_E}{[HF]_E} \Rightarrow 7.2 \times 10^{-4} = \frac{x^2}{1.0M - x}$$

	H^+	F^{1-}	HF
I	0M	0M	1.0M
C	+x	+x	-x
E	x	x	1.0M - x

Assume 5%

$1.0M - x = 1.0M$

$$7.2 \times 10^{-4} = \frac{x^2}{1.0M}$$

$$[H^+] = x = \sqrt{(7.2 \times 10^{-4})(1.0)}$$

$$[H^+] = 2.683 \times 10^{-2} M$$

$$\frac{2.683 \times 10^{-2} M}{1.0M} \cdot 100 = 2.5\%$$

$pH = -\log[H^+] =$

$-\log(2.68 \times 10^{-2} M)$

$-(-1.571)$

$pH = 1.57$

Add NaF to HF. What is new pH?
1.0M 1.0M

Assumption: We add so little volume of NaF that it will not change overall volume of solution.

$NaF \rightarrow Na^+ + F^-$ ← 1.0M NaF ($\frac{1.0 mol F^-}{1 dm^3 NaF}$)

$HF \rightleftharpoons H^+ + F^-$ $K_a = 7.2 \times 10^{-4}$

$K_a = \frac{[H^+]_e [F^-]_e}{[HF]_e}$

HF: F^- H^+ H_2O

	H^+	F^-	HF
I	0M	1.0M	1.0M
C	+x	+x	-x
E	x	1.0M+x	1.0M-x

$7.2 \times 10^{-4} = \frac{x(1.0M + x)}{1.0M - x}$

Assume 5%
$1.0M - x = 1.0M$
$1.0M + x = 1.0M$

$7.2 \times 10^{-4} = \frac{x(1)}{(1)}$

$[H^+] = x = 7.2 \times 10^{-4} M$

$pH = -\log[H^+] = -\log(7.2 \times 10^{-4} M)$
$= -(-3.142)$

$pH = 3.142$
$pH = 3.14$

Therefore, 1.0M HF pH = 1.57

① 1.0M HF + 1.0M NaF pH = 3.14

↑ pH, meaning less $[H^+]$ in the HF/NaF solution makes sense per LeChatlier's Principle (adding in F^- so system consumes F^- & H^+, so $[H^+]$↓ pH↑.

② In HF/NaF bucket, you now have significant amts of both HF & F^-

HF: HF H^+ F^- H_2O

This is called the Common Ion Effect.

This type of bucket is called a Buffer (specifically an acidic buffer).

2 Ways of making a buffer

a) Having weak acid & salt with the conjugate base of the weak acid.

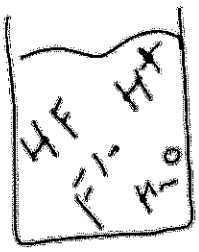
Have weak base & salt with the conjugate acid of the weak base.

b) By titrating (either w/ string acid or string base) a weak acid or weak base.

What is a buffer.

A solution that does not change its pH significantly when a strong acid or base is added.

Why?:



Remember, pH is a $\{H^+\}$ counter!!

So if add either a strong acid (H^+) or strong base (OH^-), there are species in the solution that will react with them

So Add H^+, F^- will react to form HF

Add OH^-, HF reacts to form H_2O & F^-

and as you know HF is weak so it does add H^+ to H_2O significantly.

Can you add H^+ or OH^- indefinitely?

NO, Eventually you run out of species?

(it is called exceeding the Buffer Capacity).

Using 5% Rule and the Henderson-Hasselbalch Equ (H-H) to find pH of buffered solution

2 Limitations of H-H:
1) only works for a buffer
2) 5% rule must be upheld

$HF \rightleftharpoons H^+ + F^-$

$K_a = \frac{[H^+][F^-]}{[HF]}$

1.0M HF + 1.0M NaF

ICE Box ⇒ $[HF]_E = 1.0M - x$, $[F^-]_E = 1.0M + x$, $[H^+]_E = x$

$7.2 \times 10^{-4} = \frac{[H^+][F^-]}{[HF]}$

But applying 5%, $7.2 \times 10^{-4} = \frac{[H^+](1M)}{(1M)}$

then $[H^+] \Rightarrow pH = -\log[H^+]$

$pH = 3.14$

Using Math-o-Magic

$pH = -\log[H^+]$ and $K_a = \frac{[H^+][F^-]}{[HF]}$

Sending a $-\log$ thru

$-\log K_a = -\log\left(\frac{[H^+][F^-]}{[HF]}\right)$

As you know, multiple is add with log

$-\log K_a = -\log([H^+]) + \left(-\log\frac{[F^-]}{[HF]}\right)$

Knowing $-\log(x) = p(x)$

$pK_a = pH + \left(-\log\frac{[F^-]}{[HF]}\right)$

therefore $pH = pK_a + \log\frac{[F^-]}{[HF]}$

Henderson-Hasselbalch Equ.

Example: 1.0M HF & 1.0M NaF solution

since $NaF \rightarrow Na^+ + F^-$ $[F^-] = \frac{1.0M\ NaF}{1\ mol\ NaF} \cdot \frac{1\ mol\ F^-}{1\ mol\ NaF} = 1.0M\ F^-$

$pH = pK_{a_{HF}} + \log\frac{[F^-]}{[HF]}$

$= p(7.2\times10^{-4}) + \log\frac{(1.0M)}{1.0M} \rightarrow 0$

$pH = -\log(7.2\times10^{-4})$

$pH = 3.14$

$\frac{7.2\times10^{-4}M}{1.0M} < 5\%$ ✓

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

K_a is known due to acid molecule used. Therefore need to only know $[HA] + [A^-]$ initially to determine pH of the buffer solution.

From the math of H-H equation, you can find/know 2 things

a) As long as you ratio of $\frac{[A^-]}{[HA]}$ is constant, pH of solution will be constant.

b) When $[A^-] = [HA]$, $pH = pK_a$

Adding A Strong base to an Acid Buffer - Buffer Capacity.

Remember that A Buffer has species that react with the strong BASE (OH^-) so that the H^+ does not have to react. But there is not an infinite amt available so eventually you run out (Called exceeding the Buffering Capacity)

For All system that have Strong acids or strong BASES added to them, the strong acid/BASE is react 1st & completely!!!

so you do stoichiometry to get rid of "attacking" H^+ or OH^- 1st, then make a new "bucket" with species that are left & their resulting [ ] after reaction (stoichiometry rxn).